Chapter-1

Lesson-1

- Although chemistry has its roots in alchemy, it did not fully achieve its separate status as a modern branch of science until the advent of Lavoisier's theory of combustion and of Dalton's theory of the atom. Thus as a separate modern branch of science, chemistry is just over 200 years old.
- Chemistry is the study of the properties and behavior of matter.
- Chemistry is an experimental science and concerns itself with measurement and observation.
- Chemistry is referred to as the *central science* because its study is essential to the understanding of the other branches of science.

I – The Study of Matter

Matter can be studied and classified according to its physical state (as a solid, liquid, or gas) or according to its composition (as an element, compound, or mixture).

A) States of Matter:

1) Solid – A solid has definite shape and volume (macroscopic scale). A solid's unit particles are tightly held in close and usually definite arrangements (microscopic scale). The unit particles can only vibrate in fixed positions (microscopic scale).

2) Liquid – A liquid has no definite shape but does have definite volume (macroscopic scale). A liquid's unit particles are loosely held in close yet indefinite arrangements (microscopic scale). The unit particles can move over each other (microscopic scale).

3) Gas (vapor) – A gas has neither a definite shape nor volume (macroscopic scale). A gas's unit particles are not held to one another (microscopic scale). A gas's unit particles are widely separated and are constantly moving around at high speeds (microscopic scale).

B) Composition of Matter:

1) Substance – A substance is a *pure form of matter* that has distinct set of properties and a composition that does not vary from sample to sample. Substances can be further classified as ***elements*** or ***compounds***.

   a) *Elements* cannot easily be decomposed into simpler substances. Elements are composed of a single kind of atom.

   Examples: carbon, hydrogen, oxygen, sodium, iron, and silicon

   b) *Compounds* can easily be decomposed into simpler substances. A compound is composed of two or more elements united chemically in definite proportions. This is known as the law of constant composition (or the law of definite proportions). Compounds have properties that are different from the elements that compose them.

   Examples: water, glucose, sodium chloride, and carbon dioxide

2) Mixture –A mixture is an *impure form of matter* composed of two or more substances (elements and/or compounds) in an indefinite proportion. The components in a mixture usually retain their own properties. Mixtures can be further classified as being either ***homogeneous*** or ***heterogeneous***.

   a) *Homogeneous Mixtures* have a uniform composition throughout the sample. ***Solutions*** are the only class of mixture that is homogeneous at both the macroscopic and microscopic scale.

   Examples: air, brine, and bronze

   b) *Heterogeneous Mixtures* have a varied composition throughout the sample. ***Suspensions***, ***colloids****, ***gels***, ***sols***, and ***foams***, are some classes of heterogeneous mixtures.

   Examples: dirt, rocks, and wood

***Note:** Colloids are mixtures that are classified as heterogeneous at the microscopic scale but that could be classified as homogeneous at the macroscopic scale. The uniformity of mixture can depend on the scale at which observations are made.

C) Properties of Matter:

1) Physical Property – A property that can be determined without changing the identity and composition of the sample being observed.

   Examples: color, odor, taste, mass, volume, density, melting point, boiling point, solubility and hardness

2) Chemical Property – A property that describes the manner in which a substance may or may not change to form new substances.

   Examples: flammability and corrosiveness to acid

3) Intensive Property – A property that is true regardless of the quantity of matter that is present.

   Examples: density, melting point, flammability, and molar solubility

4) Extensive Property – A property that does depend on the quantity of matter that is present.

   Examples: mass, volume, temperature, and length

D) Changes in Matter:

1) Physical Change – A change (reaction) that alters the physical appearance but not the chemical composition of a sample of matter.

   Examples: all phase changes, most dissolution, and breaking

2) Chemical Change – A change (reaction) that alters the chemical composition of a sample of matter.

   Examples: combustion, decomposition, and corrosion

3) Ambiguous Change – A change (reaction) that can be argued as either physical or chemical.

   Examples: dissolution or precipitation of an ionic substance**

   ****Note:** Some would say that the dissolution (into water) and precipitation (from water) of an ionic substance is simply a physical change since no new substances were formed. Others would say these changes are chemical since ionic bonds were broken/formed and ion-dipole forces were formed/broken.

II – The Separation of Mixtures

Most samples of matter in nature are mixtures. Mixtures can be separated by various techniques due to differences in properties of the components in the mixture.

A) Filtration: *(due to differences in particle size)*

Used to separate an insoluble solid from a liquid or a solution.

B) Distillation: *(due to differences in boiling point)*

Used to separate a liquid from a dissolved solid.
Used to separate a liquid (A) from a liquid (B).
"A" must have a lower boiling point than "B" by at least 5°C

C) Crystallization: *(due to differences in solubility)*

Used to separate a solid (A) from a solid (B).
"A" must be less soluble in the chosen solvent and the quantity of "B" in the mixture must be small.

D) Chromatography: *(due to differences in polarity)*

Used to separate two substances, (A) and (B), that have differing attractions to a solvent due to their polarities.
"A" and "B" must be attracted differently from each other to the stationary phase (the paper/the material in the column) and the mobile phase (the solvent) being used.

1) Paper Chromatography

2) Column chromatography

E) Extraction: *(due to differences in polarity)*

Used to separate a dissolved substance (A) from a dissolved substance (B) by using a new solvent that is immiscible with the original solvent.
"A" must be highly soluble in the new solvent but "B" must be highly insoluble in the new solvent.

Chapter-1

Lesson-2

III – Units of Measurement

The preferred metric units of measurement are called SI units. There are seven basic SI units:

kilogram, meter, second, Kelvin, mole, ampere, and candela.
(We will use all of these units this year except for the candela.)

A) Length:

- The meter (m) is the SI base unit of length.
- Prefixes are used to change the magnitude of the unit:

mega,	**kilo,**	**deci,**	**centi,**	**milli,**	**micro,**	**nano,**	**pico**
Mm,	**km,**	**dm,**	**cm,**	**mm,**	**µm,**	**nm**	**pm**
10^6m	**10^3m,**	**10^{-1}m,**	**10^{-2}m,**	**10^{-3}m,**	**10^{-6}m,**	**10^{-9}m,**	**10^{-12}m**

B) Volume:

- The cubic meter (m^3) is the SI unit of volume.
- 1 m^3 = 1000 dm^3 (*1m = 10 dm*)
- 1 dm^3 = 1 L
- 1 L = 1000 mL
- 1 mL = 1 cm^3
- There are several instruments to measure volume:
  - graduated cylinder, pipette, buret, volumetric flask

C) Mass:

- The kilogram is the SI base unit of mass.
- 1 kg = 1000 g
- 1 g = 1000 mg
- Mass vs Weight
  - weight = mass × gravity
  - a balance can measure mass because…

    g×(mass of sample) = g×(known mass)

D) Temperature:

- Celsius is the temperature unit for science.
- Many calculations require Kelvin temperature

$$K = C + 273$$

- We live our everyday lives (USA) using Fahrenheit temperature.

$$F = 1.8C + 32$$

Q1: $37.0^{o}C = ?\ ^{o}F$

A1: $\boxed{98.6^{o}F}$

Q2: $37.0^{o}C = ?\ K$

A2: $\boxed{310.0\ K}$

IV - Experimental Error and Significant Figures

All measurements contain uncertainty (error) which can be attributed to the instrument being used as well as the individual using the instrument. The instrument being used will determine those digits in the measurement that are significant.

A) Uncertainty in Addition and Subtraction:

Q3: 32.11 g + 1.4 g = ?

A3: 33.51 g $\rightarrow \boxed{33.5\ g}$

Q4: 11 mL - 5.25 mL = ?

A4: 5.75 mL $\rightarrow \boxed{6\ mL}$

B) Uncertainty in Multiplication and Division:

Q5: $4.18\ J/g^{o}C \times 5.2^{o}C = ?$

A5: 21.736 $\rightarrow \boxed{22\ J/g}$

Q6: $16.004\ g \div 5.9\ mL = ?$

A6: 2.712542373 $\rightarrow \boxed{2.7\ g/mL}$

C) Exact Numbers or Values by Definition:

Exact numbers (counting numbers) or values by definition do not affect significant figures.

Q7: If 1 atom of carbon has a mass of 12.01 μ, then the mass of 5 atoms is…

A7: $5 \times 12.01\ \mu = \boxed{60.05\ \mu}$ ← note that the answer to the same place as the original value.

Q8: If by definition 1 calorie is the amount of heat needed to raise the temperature of 1 g of water by 1°C, then the amount of heat needed to raise 5.736 g H_2O by 6.5°C is...

A8: $5.736\ g \times 1\ cal/g^oC \times 6.5^oC = 37.284 \rightarrow \boxed{37\ cal}$ ← two sig. figs.

V –Conversion Factors & Dimensional Analysis

Dimensional analysis is a technique by which a given value is multiplied by a conversion factor (a unit ratio) such that the final answer contains the desired units.

A) Single Unit Conversions:

Q9: How many moles of carbon atoms are contained in 100.0 g of carbon?

A9: $$100.0\ g\cancel{C} \times \frac{1\ mole}{12.011\ g\cancel{C}} = \boxed{8.326\ moles\ C}$$

B) Multiple Unit Conversions:

Q10: What is the speed of light (3.00×10^8 m/s) in miles per hour?

A10: $$3.00\times10^8\ \frac{\cancel{m}}{\cancel{s}} \times \frac{3600\ \cancel{s}}{1\ hr} \times \frac{1mi}{1600\ \cancel{m}} = \boxed{6.8\times10^8\ mi/hr}$$

Q11: If 1.000 mol of gas at STP has a volume of 22.4 dm^3 then determine its volume in ft^3.

A11: $$22.4\ d\cancel{m}^3 \times \frac{(10\ c\cancel{m})^3}{(1\ d\cancel{m})^3} \times \frac{(1.00\ \cancel{in})^3}{(2.54\ c\cancel{m})^3} \times \frac{(1\ \cancel{ft})^3}{(12\ \cancel{in})^3} = \boxed{0.791\ ft^3}$$

VI – Rules for Rounding

In most calculations, you will need to round numbers to obtain the correct number of significant figures. The following rules should be applied when rounding

1) In a series of calculations, carry the extra digits through to the final result, THEN round.

2) If the digit to be removed…

   a) …is less than 5, the preceding digit remains the same.

   Example: 1.33 cm (2 sig. figs.) rounds to 1.3 cm

   b) …is equal to or more than 5, the preceding digit is increased by 1.

   Example: 4.348 g (3 sig. figs.) rounds to 4.35 g

Q12: A penny made before 1982 has a mass of 3.027 g and a density of 8.9 g/cm^3. Calculate its volume in cm^3

A12: d =m/V $\rightarrow$ V = m/d $\rightarrow$ V = 3.027 g ÷ 8.9 g/cm^3

$\boxed{V = 0.34\ \text{cm}^3}$

Q13: Assuming all pre-1982 pennies have a mass of 3.027 g, how many pennies would be in a pound? (454 g = 1.00 lb)

A13: $454\ \text{g} \times \dfrac{1\ \text{penny}}{3.027\ \text{g}} = \boxed{150.\ \text{pennies}}$

Q14: If a penny's diameter is 1.9 cm, then what is its thickness?

A14: $V = \pi r^2 h$

$0.34\ \text{cm}^3 = (\pi)(1.9/2\ \text{cm})^2(x)$

$\boxed{x = 0.12\ \text{cm}}$